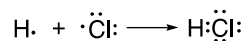


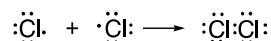
Chapter 4 Covalent Compounds

Solutions to In-Chapter Problems

- 4.1 H is surrounded by two electrons, giving it the noble gas configuration of He. Cl is surrounded by eight electrons, giving it the noble gas configuration of Ar.



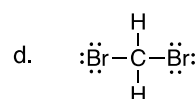
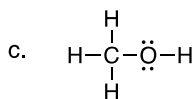
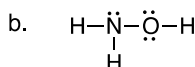
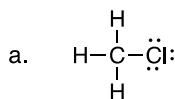
- 4.2 This arrangement gives each Cl an octet and the electronic configuration of Ar.



- 4.3 Atoms with one, two, or three valence electrons form one, two, or three bonds, respectively. Atoms with four or more valence electrons form enough bonds to give an octet.

- | | | |
|--|--|--|
| a. F forms one bond.
(8 - 7 valence e^- = 1 bond) | c. Br forms one bond.
(8 - 7 valence e^- = 1 bond) | e. P forms three bonds.
(8 - 5 valence e^- = 3 bonds) |
| b. Si forms four bonds.
(8 - 4 valence e^- = 4 bonds) | d. O forms two bonds.
(8 - 6 valence e^- = 2 bonds) | f. S forms two bonds.
(8 - 6 valence e^- = 2 bonds) |

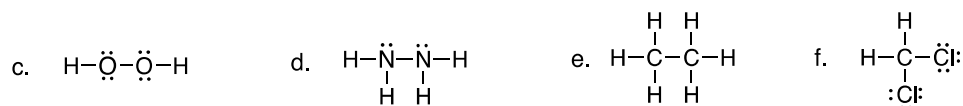
4.4



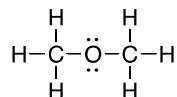
- 4.5 Ionic bonding is observed in CaO since Ca is a metal and readily transfers electrons to the nonmetal oxygen. Covalent bonding is observed in CO_2 since carbon is a nonmetal and does not readily transfer electrons.

4.6

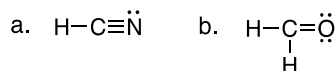
<p>Step [1] a. Arrange the atoms.</p> <p style="text-align: center;">H Br</p>	<p>Step [2] Count the electrons.</p> <p style="text-align: center;">1 Br x 7 e^- = 7 e^- 1 H x 1 e^- = 1 e^- <hr style="width: 50%; margin: 0 auto;"/>8 e^- total</p>	<p>Step [3] Add the bonds and lone pairs.</p> <p>Add bonds first... ...then lone pairs.</p> <p style="text-align: center;">$\text{H}-\text{Br} \quad \text{---} \quad \text{H}-\ddot{\text{Br}}\text{:}$ only 2 e^- used 8 e^- used</p>
<p>Step [1] b. Arrange the atoms.</p> <p style="text-align: center;">H H C F H</p> <p>• four atoms around C</p>	<p>Step [2] Count the electrons.</p> <p style="text-align: center;">1 C x 4 e^- = 4 e^- 1 F x 7 e^- = 7 e^- 3 H x 1 e^- = 3 e^- <hr style="width: 50%; margin: 0 auto;"/>14 e^- total</p>	<p>Step [3] Add the bonds and lone pairs.</p> <p>Add bonds first... ...then lone pairs.</p> <p style="text-align: center;">$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{F} \\ \\ \text{H} \end{array} \quad \text{---} \quad \begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\ddot{\text{F}}\text{:} \\ \\ \text{H} \end{array}$ no octet 14 e^- used only 8 e^- used</p>



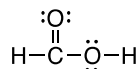
4.7 Use the same steps as in Answer 4.6 to draw the Lewis structure.



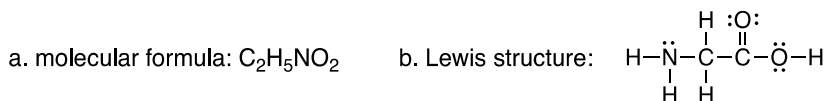
4.8 After placing all electrons in bonds and lone pairs, use a lone pair to form a multiple bond if an atom does not have an octet. Follow the stepwise procedure in Example 4.3.



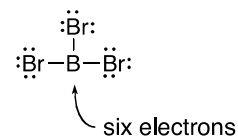
4.9 A valid Lewis structure for formic acid contains a carbon–oxygen double bond. The C atom has four bonds and each O atom has two bonds and two lone pairs.



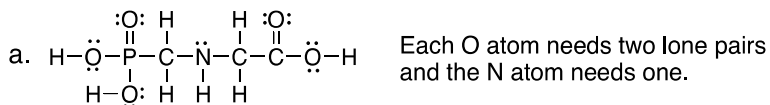
4.10



4.11 Boron is surrounded by only six electrons, so the Lewis structure for BBr_3 does not follow the octet rule.



4.12

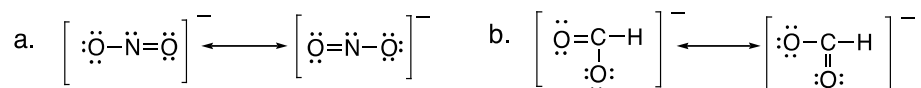


- b. There are 10 electrons around phosphorus.
c. Phosphorus and hydrogen do not follow the octet rule.

4.13 Resonance structures are two or more Lewis structures having the same arrangement of atoms but a different arrangement of electrons. Draw the second resonance structure as in Sample Problem 4.5 in the text, or Example 4.4. The two resonance structures in each part differ in the location of a double bond and a lone pair.

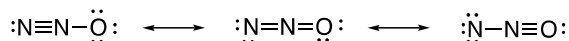


4.14



4.15

Three resonance structures for nitrous oxide:



4.16 Name the compounds using the following two-step method:

Step [1] Name the first nonmetal by its element name and the second using the suffix *-ide*.

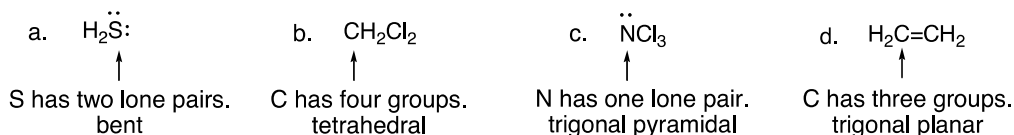
Step [2] Add prefixes to show the number of atoms of each element.

a. CS₂ = carbon disulfidec. PCl₅ = phosphorus pentachlorideb. SO₂ = sulfur dioxided. BF₃ = boron trifluoride

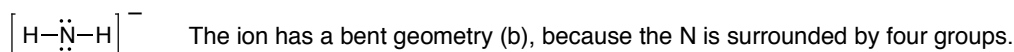
4.17

a. silicon dioxide = SiO₂c. sulfur trioxide = SO₃b. phosphorus trichloride = PCl₃d. dinitrogen trioxide = N₂O₃

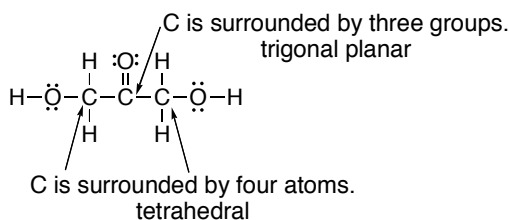
4.18



4.19



4.20 Dihydroxyacetone has (a) two tetrahedral carbons and (b) one trigonal planar carbon.

4.21 Electronegativity *increases* left to right across a row of the periodic table as the nuclear charge increases (excluding the noble gases).Electronegativity *decreases* down a column of the periodic table as the atomic radius increases, pushing the valence electrons farther from the nucleus.

a. Na < Li < H

b. Be < C < O

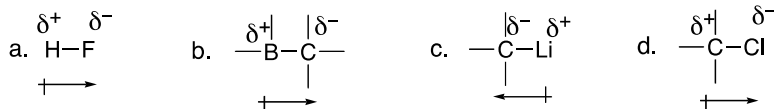
c. I < Cl < F

d. B < N < O

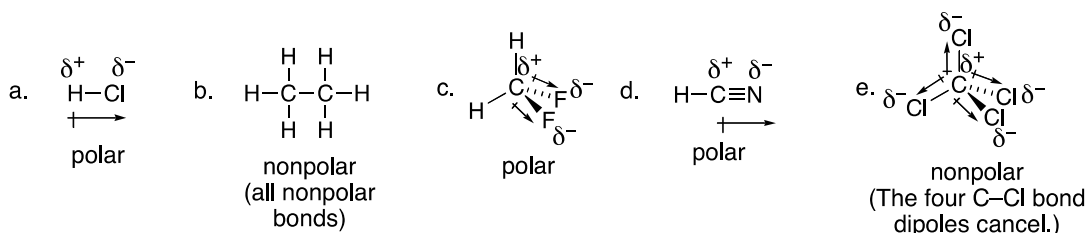
4.22 Calculate the electronegativity difference between the two atoms and use the following rules: less than 0.5 (nonpolar); 0.5–1.9 (polar covalent); greater than 1.9 (ionic).

- a. HF = (4.0 – 2.1) = 1.9 = polar covalent d. ClF = (4.0 – 3.0) = 1.0 = polar covalent
 b. MgO = (3.5 – 1.2) = 2.3 = ionic e. H₂O = (3.5 – 2.1) = 1.4 = polar covalent
 c. F₂ = (4.0 – 4.0) = 0 = nonpolar f. NH₃ = (3.0 – 2.1) = 0.9 = polar covalent

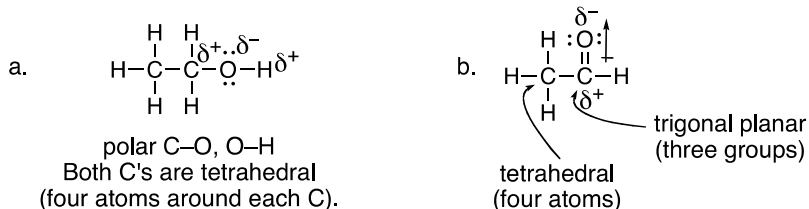
4.23 The head of the bond dipole arrow points towards the more electronegative atom.



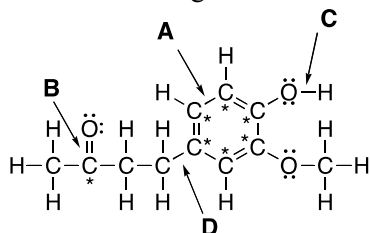
4.24 To determine the overall polarity of a molecule, identify the polar bonds, determine the shape around individual atoms, and then decide if the individual bond dipoles cancel or reinforce.



4.25 All C–H and C–C bonds are nonpolar.



4.26 a. Lewis structure of zingerone:



Each O atom has two lone pairs, so there is a total of six lone pairs (containing 12 electrons) in zingerone.

- b. There are seven trigonal planar C's, labeled with (*).
 c. Bonds **A** and **D** are between two carbon atoms so the bonds are nonpolar. Since the electronegativity of O (3.5) is much greater than the electronegativity of C (2.5) and H (2.1), the C–O bond (**B**) and O–H bond (**C**) are both polar.

Solutions to End-of-Chapter Problems

4.27 In covalent bonding, atoms share electrons to attain the electronic configuration of the noble gas closest to them in the periodic table. In ionic bonding, one atom donates electrons to the other atom.

- a. LiCl: ionic; the metal Li donates electrons to the nonmetal chlorine.
 HCl: covalent; H and Cl share electrons because both atoms are nonmetals and the electronegativity difference is not large enough for electron transfer to occur.
- b. KBr: ionic; the metal K donates electrons to the nonmetal bromine.
 HBr: covalent; H and Br share electrons because both atoms are nonmetals and the electronegativity difference is not large enough for electron transfer to occur.

4.28 In covalent bonding, atoms share electrons to attain the electronic configuration of the noble gas closest to them in the periodic table. In ionic bonding, one atom donates electrons to the other atom.

- a. BeH₂: covalent; Be is a metal and H is a nonmetal, but the electronegativity difference between Be and H (2.1-1.5 = 0.6) is not large enough for electron transfer to occur.
 BeCl₂: covalent; Be is a metal and Cl is a nonmetal, but the electronegativity difference between Be and Cl (3.0 - 1.5 = 1.5) is not large enough for electron transfer to occur.
- b. Na₃N: ionic; the metal Na donates electrons to the nonmetal chlorine.
 NH₃: covalent; N and H share electrons because both atoms are nonmetals and the electronegativity difference is not large enough for electron transfer to occur.

4.29 Atoms with one, two, or three valence electrons form one, two, or three bonds, respectively. Atoms with four or more valence electrons form enough bonds to give an octet.

- a. C = 4 bonds, 0 lone pairs
 b. Se = 2 bonds, 2 lone pairs
 c. I = 1 bond, 3 lone pairs
 d. P = 3 bonds, 1 lone pair

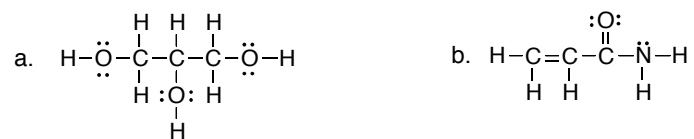
4.30 Atoms with one, two, or three valence electrons form one, two, or three bonds, respectively. Atoms with four or more valence electrons form enough bonds to give an octet.

- a. O = 2 bonds, 2 lone pairs
 b. Si = 4 bonds, 0 lone pairs
 c. Ge = 4 bonds, 0 lone pairs
 d. B = 3 bonds, 0 lone pairs

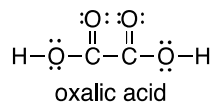
4.31 Each O and S atom must get two lone pairs and each N atom must get one lone pair.



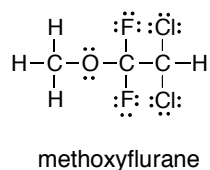
4.32 Each O atom must get two lone pairs and the N atom must get one lone pair.



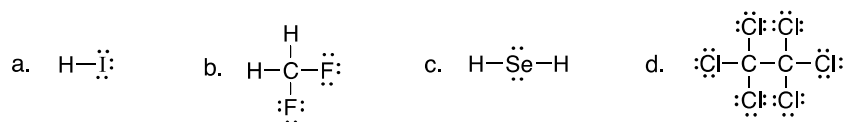
4.33 Use the common element colors (black = C, red = O, gray = H), and give each O two lone pairs.



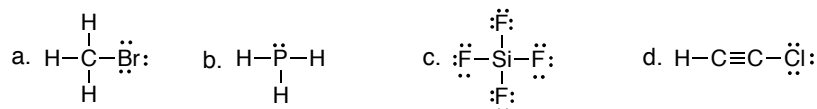
4.34 Use the common element colors (black = C, red = O, gray = H, yellow = F, green = chlorine), and give each O two lone pairs and each F and Cl three lone pairs.



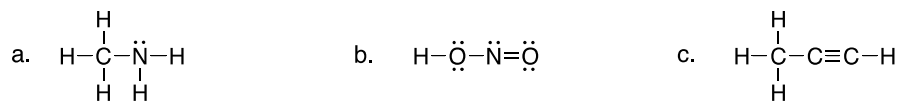
4.35 Follow the steps in Examples 4.2 and 4.3 to draw the Lewis structures.



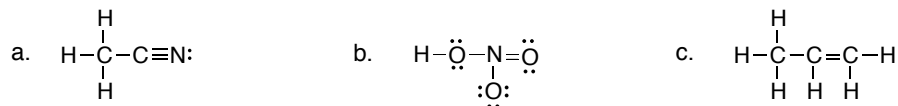
4.36 Follow the steps in Examples 4.2 and 4.3 to draw the Lewis structures.



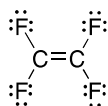
4.37 Follow the steps in Examples 4.2 and 4.3.



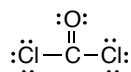
4.38 Follow the steps in Examples 4.2 and 4.3.



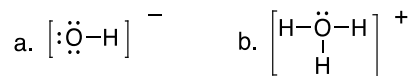
4.39 Follow the steps in Example 4.3.



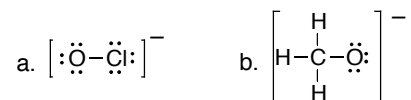
4.40 Follow the steps in Example 4.3.



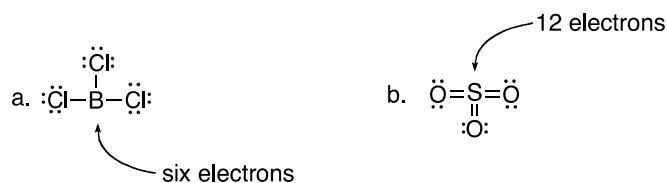
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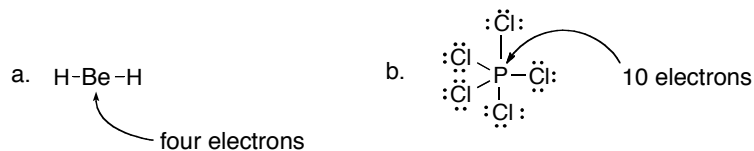
4.42



4.43



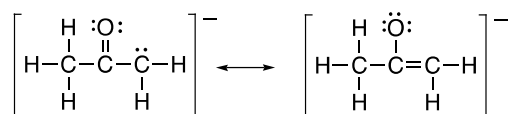
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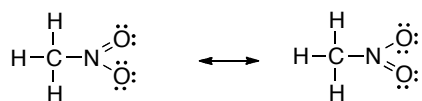
4.45 A resonance structure is one representation of electron arrangement with a given placement of atoms. A resonance hybrid is a composite of two or more resonance structures.

4.46 Having two resonance structures for a molecule stabilizes it because the lone pairs and electron pairs in multiple bonds are spread out over a larger volume of space.

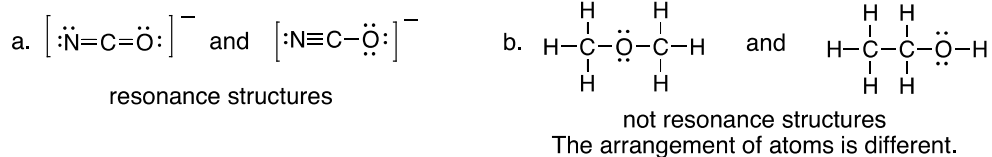
4.47 Use Example 4.4. In this ion, the location of a double bond and a lone pair is different in the two resonance structures.



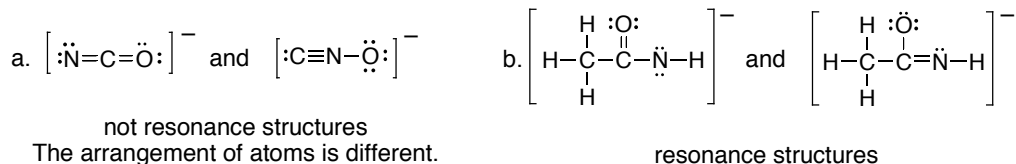
4.48 Use Example 4.4. In this molecule, the location of a double bond and a lone pair is different in the two resonance structures.



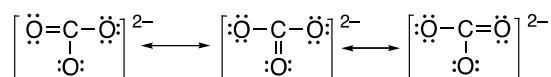
4.49 Resonance structures must have the same arrangement of atoms.



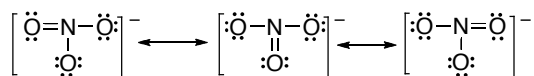
4.50 Resonance structures must have the same arrangement of atoms.



4.51 Three resonance structures for the carbonate anion:



4.52 Three resonance structures for the nitrate anion:



4.53 Name each compound as in Answer 4.16.

- a. PBr_3 = phosphorus tribromide
 b. SO_3 = sulfur trioxide
 c. NCl_3 = nitrogen trichloride
 d. P_2S_5 = diphosphorus pentasulfide

4.54 Name each compound as in Answer 4.16.

- a. SF_6 = sulfur hexafluoride
 b. CBr_4 = carbon tetrabromide
 c. N_2O = dinitrogen monoxide
 d. P_4O_{10} = tetraphosphorus decoxide

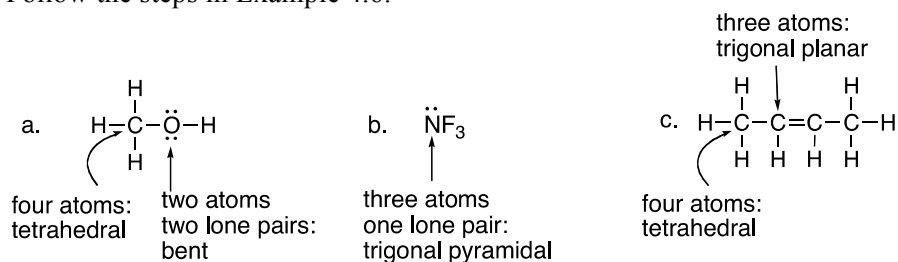
4.55 Work backwards to write the formula that corresponds to each name.

- a. selenium dioxide = SeO_2
 b. carbon tetrachloride = CCl_4
 c. dinitrogen pentoxide = N_2O_5

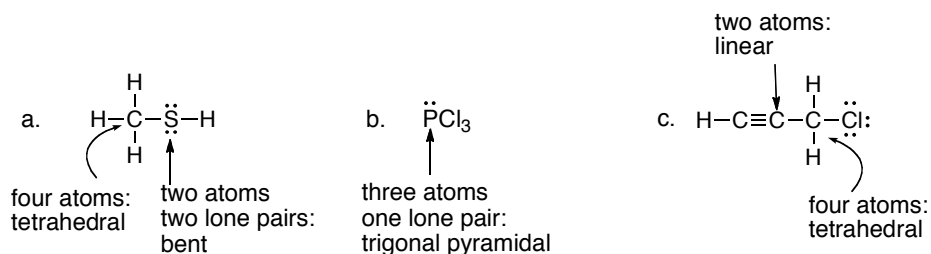
4.56 Work backwards to write the formula that corresponds to each name.

- a. silicon tetrafluoride = SiF_4
 b. nitrogen monoxide = NO
 c. phosphorus triiodide = PI_3

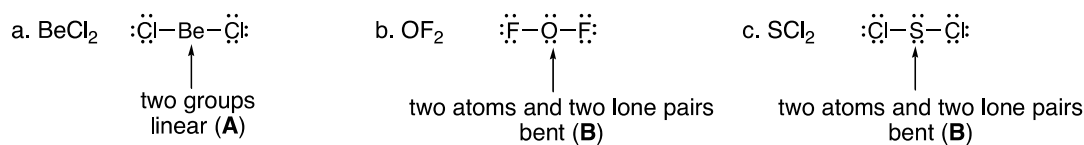
4.57 Follow the steps in Example 4.6.



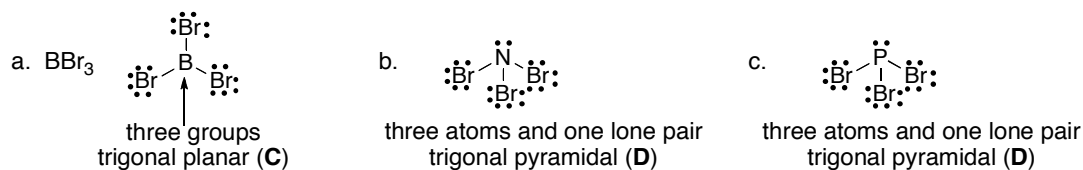
4.58 Follow the steps in Example 4.6.



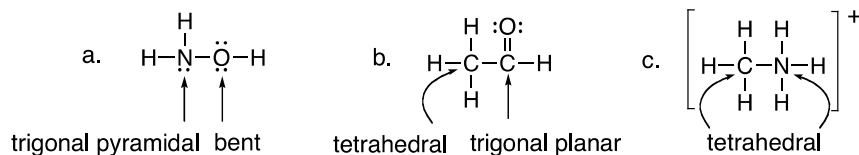
4.59 First, draw a Lewis structure; then count groups around the central atom to determine shape. With two groups, an atom is linear. With two atoms and two lone pairs, an atom has a bent shape.



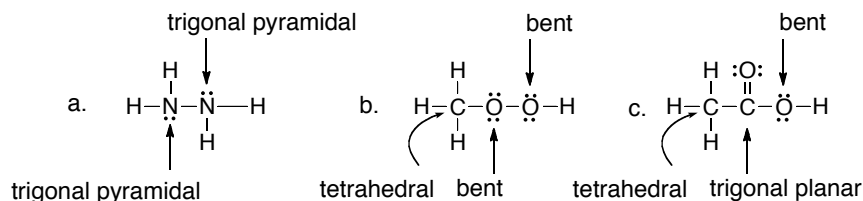
4.60 First, draw a Lewis structure; then count groups around the central atom to determine shape. With three groups, an atom is trigonal planar. With three atoms and one lone pair, an atom has a trigonal pyramidal shape.



4.61 Follow the steps in Example 4.6.



4.62 Follow the steps in Example 4.6.



4.63

- a. [1] A linear geometry means two groups, so the dark red atom has no lone pairs. [2] BeCl_2
 b. [1] A bent geometry generally means four groups, so the dark red atom has two lone pairs. [2] H_2O

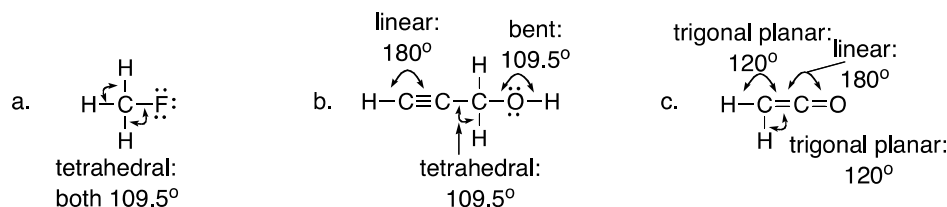
4.64

- a. [1] A trigonal planar geometry means three groups, so the dark red atom has no lone pairs. [2] BCl_3
 b. [1] A trigonal pyramidal geometry generally means four groups, so the dark red atom has one lone pair. [2] NH_3

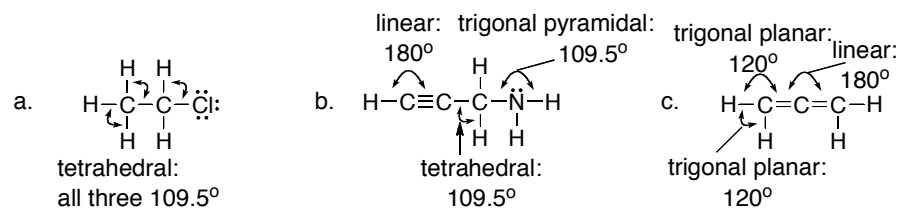
4.65 BCl_3 is trigonal planar because it has three Cl's bonded to B but no lone pairs. NCl_3 is trigonal pyramidal because it has three Cl's around N as well as a lone pair.

4.66 H_2O is bent because it has two H's bonded to O and two lone pairs. H_3O^+ is trigonal pyramidal because it has three H's bonded to O and one lone pair.

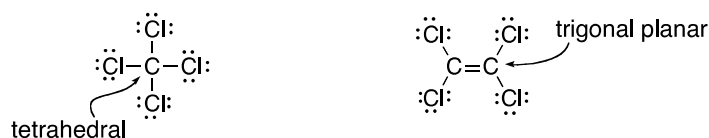
4.67



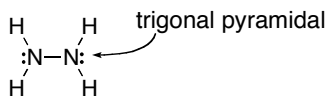
4.68



4.69 CCl_4 is tetrahedral because carbon has four groups around it. Each C in C_2Cl_4 has trigonal planar geometry because each carbon has three groups around it.



- 4.70 Each N in N_2H_4 has trigonal pyramidal geometry because each nitrogen has three atoms (i.e., two H's and one N) and one lone pair around it.



- 4.71 Electronegativity *increases* left to right across a row of the periodic table as the nuclear charge increases (excluding the noble gases).
Electronegativity *decreases* down a column of the periodic table as the atomic radius increases, pushing the valence electrons farther from the nucleus.

- a. $\text{Se} < \text{S} < \text{O}$ b. $\text{Na} < \text{P} < \text{Cl}$ c. $\text{S} < \text{Cl} < \text{F}$ d. $\text{P} < \text{N} < \text{O}$

- 4.72 Electronegativity *increases* left to right across a row of the periodic table as the nuclear charge increases (excluding the noble gases).
Electronegativity *decreases* down a column of the periodic table as the atomic radius increases, pushing the valence electrons farther from the nucleus.

- a. $\text{Si} < \text{P} < \text{S}$ b. $\text{Ca} < \text{Mg} < \text{Be}$ c. $\text{Se} < \text{Br} < \text{Cl}$ d. $\text{Na} < \text{Li} < \text{Be}$

- 4.73 In covalent bonding, atoms share electrons to attain the electronic configuration of the noble gas closest to them in the periodic table. In ionic bonding, one atom donates electrons to the other atom. When the electronegativity difference between the two atoms in a bond is greater than 1.9, the bond is ionic.

- a. hydrogen and bromine = polar covalent
($2.8 - 2.1 = 0.7$) c. sodium and sulfur = polar covalent
($2.5 - 0.9 = 1.6$)
b. nitrogen and carbon = polar covalent d. lithium and oxygen = ionic
($3.0 - 2.5 = 0.5$) ($3.5 - 1.0 = 2.5$)

- 4.74 In covalent bonding, atoms share electrons to attain the electronic configuration of the noble gas closest to them in the periodic table. In ionic bonding, one atom donates electrons to the other atom. When the electronegativity difference between the two atoms in a bond is greater than 1.9, the bond is ionic.

- a. nitrogen and oxygen = polar covalent
($3.5 - 3.0 = 0.5$) c. sulfur and chlorine = polar covalent
($3.0 - 2.5 = 0.5$)
b. oxygen and hydrogen = polar covalent d. sodium and chlorine = ionic
($3.5 - 2.1 = 1.4$) ($3.0 - 0.9 = 2.1$)

- 4.75 Calculate the electronegativity difference between the atoms and carbon and use the following rules: less than 0.5 (nonpolar); 0.5–1.9 (polar covalent); greater than 1.9 (ionic).

- a. $\text{C} = (2.5 - 2.5) = 0 = \text{nonpolar}$ d. $\text{Cl} = (3.0 - 2.5) = 0.5 = \text{polar}$
b. $\text{O} = (3.5 - 2.5) = 1.0 = \text{polar}$ e. $\text{H} = (2.5 - 2.1) = 0.4 = \text{nonpolar}$
c. $\text{Li} = (2.5 - 1.0) = 1.5 = \text{polar}$

4.76 Calculate the electronegativity difference between the atoms and fluorine and use the following rules: less than 0.5 (nonpolar); 0.5–1.9 (polar covalent); greater than 1.9 (ionic).

- a. $H = (4.0 - 2.1) = 1.9 = \text{polar}$ d. $Li = (4.0 - 1.0) = 3.0 = \text{ionic}$
 b. $F = (4.0 - 4.0) = 0 = \text{nonpolar}$ e. $S = (4.0 - 2.5) = 1.5 = \text{polar}$
 c. $C = (4.0 - 2.5) = 1.5 = \text{polar}$

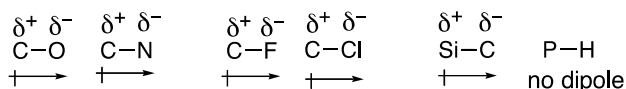
4.77 Calculate the electronegativity difference between the atoms to determine which bond is more polar.

- a. $C-O = (3.5 - 2.5) = 1.0 = \text{more polar}$ c. $Si-C = (2.5 - 1.8) = 0.7 = \text{more polar}$
 $C-N = (3.5 - 3.0) = 0.5$ $P-H = (2.1 - 2.1) = 0$
 b. $C-F = (4.0 - 2.5) = 1.5 = \text{more polar}$
 $C-Cl = (3.0 - 2.5) = 0.5$

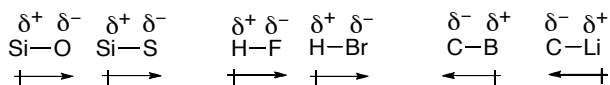
4.78 Calculate the electronegativity difference between the atoms to determine which bond is more polar.

- a. $Si-O = (3.5 - 1.8) = 1.7 = \text{more polar}$ c. $C-B = (2.5 - 2.0) = 0.5$
 $Si-S = (2.5 - 1.8) = 0.7$ $C-Li = (2.5 - 1.0) = 1.5 = \text{more polar}$
 b. $H-F = (4.0 - 2.1) = 1.9 = \text{more polar}$
 $H-Br = (2.8 - 2.1) = 0.7$

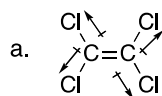
4.79 The δ^+ is placed next to the *less* electronegative atom, whereas the δ^- is placed next to the *more* electronegative atom.



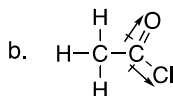
4.80 The δ^+ is placed next to the less electronegative atom, and the δ^- is placed next to the more electronegative atom.



4.81

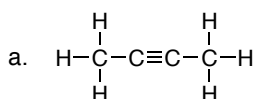


polar bonds
nonpolar compound
Bond dipoles cancel.

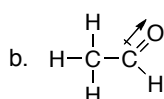


polar bonds
polar compound
Bond dipoles do not cancel.

4.82

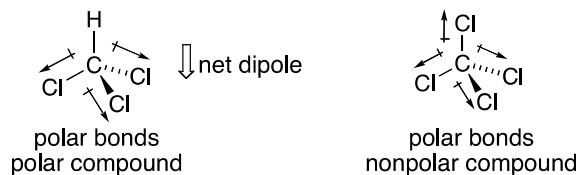


nonpolar bonds
nonpolar compound



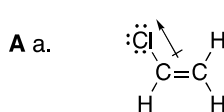
polar bond
polar compound

- 4.83 CHCl_3 has a net dipole, making it polar. CCl_4 has four polar bonds but no net dipole because the four bond dipoles cancel one another.

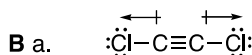


- 4.84 H_2O has two polar O–H bonds. The molecule has a net dipole, making it polar. The H–S bonds in H_2S are nonpolar because the electronegativity difference between H and S is only 0.4. As a result, H_2S is nonpolar.

4.85

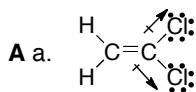


- b. one polar C–Cl bond
 c. polar molecule
 (Only one polar bond dipole makes a polar molecule.)

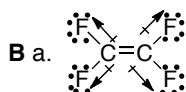


- b. two polar C–Cl bonds
 c. nonpolar molecule
 (Two bond dipoles cancel.)

4.86



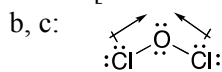
- b. two polar C–Cl bonds
 c. polar molecule
 bond dipoles do not cancel



- b. four polar C–F bonds
 c. nonpolar molecule
 bond dipoles cancel

4.87

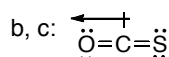
- a. 20 valence electrons
 [1 O atom (6 valence electrons) + 2 Cl atoms (7 valence electrons each)]



- d. Cl_2O has a bent shape because O is surrounded by two atoms and two lone pairs.
 e. The compound is polar because the two bond dipoles do not cancel.

4.88

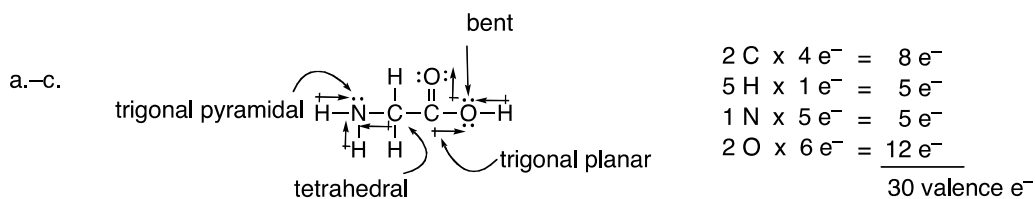
- a. 16 valence electrons [1 O atom (6 valence electrons) + 1 C atom (4 valence electrons) + 1 S atom (6 valence electrons)]



- d. OCS has a linear shape because C is surrounded by two atoms and no lone pairs.

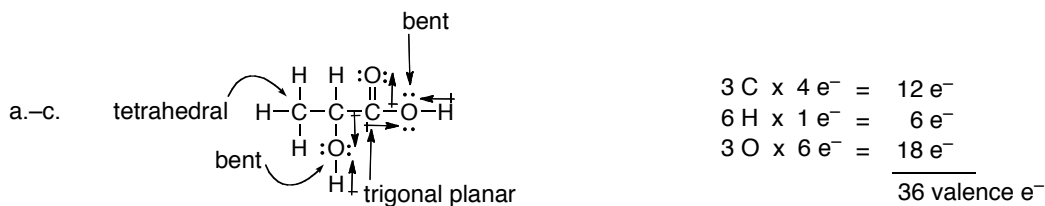
- e. Yes, the compound is polar because there is only one bond dipole. The bond between carbon and sulfur is nonpolar because the electronegativity values of both C and S are 2.5.

4.89



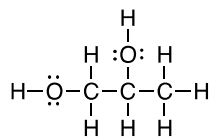
d. Glycine is a polar molecule because the many bond dipoles do not cancel.

4.90



d. Lactic acid is a polar molecule because the many bond dipoles do not cancel.

4.91 a. Lewis structure:



b. Both C–O and O–H bonds are polar.

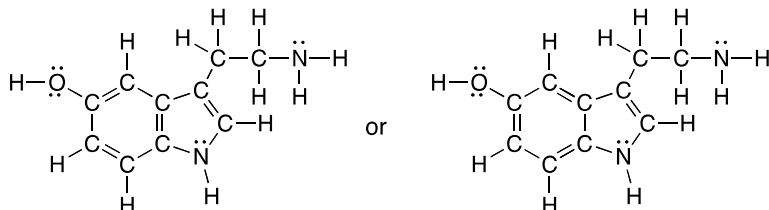
c. All C's are surrounded by four atoms, making them tetrahedral. Both O's are surrounded by two atoms and two lone pairs, giving them a bent shape.

4.92

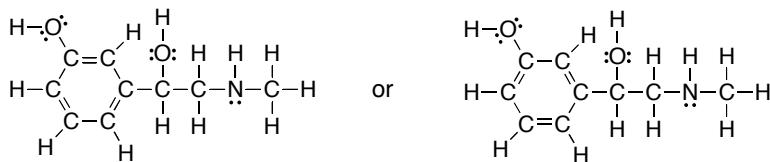


polar bonds
polar compound
F more electronegative than Cl

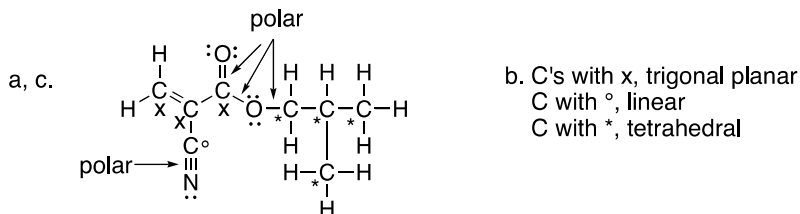
4.93 First, give the O atom two lone pairs and each N atom one lone pair. Then add double bonds to the C's that do not have four bonds.



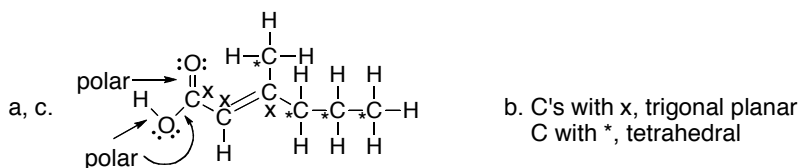
- 4.94 First, give the two O atoms two lone pairs and the N atom one lone pair. Then add double bonds to the C's that do not have four bonds.



4.95

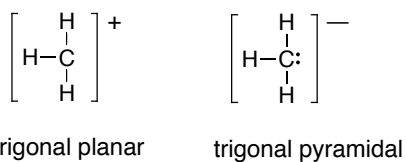


4.96

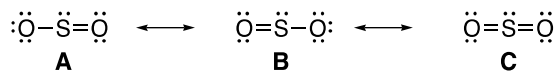


- 4.97
- The predicted shape around each carbon is tetrahedral because each C is surrounded by four atoms.
 - The interior angle of a triangle with three equal sides is 60° .
 - The 60° bond angle deviates greatly from the theoretical tetrahedral bond angle (109.5°), making the molecule unstable.

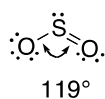
4.98



4.99



Structures **A** and **B** follow the octet rule. Since the S atom in **C** has 10 electrons, this structure violates the octet rule. Because each structure has an S atom surrounded by two atoms and a lone pair, the three groups must be $\sim 120^\circ$ from each other. For example, **A** could be drawn as:



4.100

